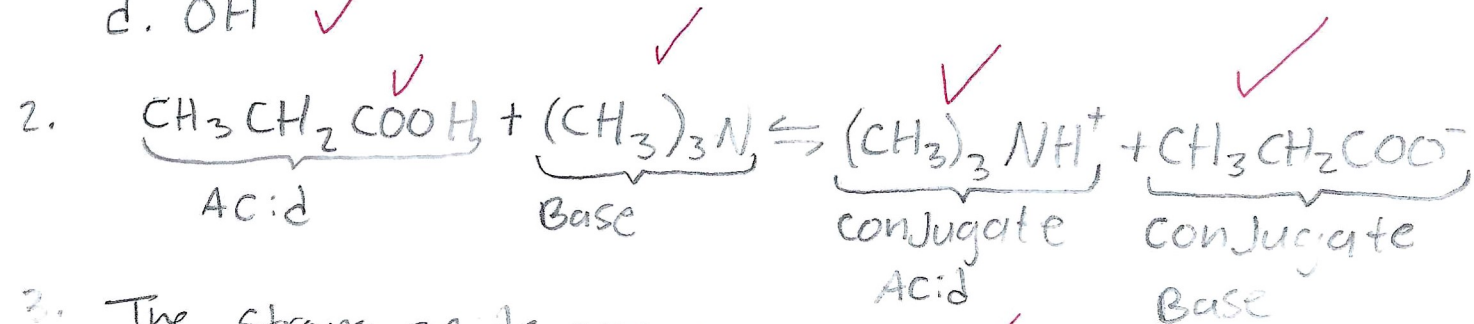


$$\frac{26}{26} = 100\%$$



3. The strong acids are c and g ✓
 The strong bases are f and b ✓

4. The difference between the Lewis and Brønsted-Lowry definition of acids is that the Lewis definition says lone pairs of electrons are accepted whereas in the Brønsted-Lowry definition, H^+ are accepted ✓

5. Yes, it is possible for a Lewis acid to also be a Brønsted-Lowry ✓
 6. F^- is the base and BeF_2 is the Acid ✓

7. pH of 0.040 M solution of HNO_3 ? $\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$ ✓

$$\left(\frac{0.040 \text{ moles of } \text{HNO}_3}{1 \text{ liter}} \right) \left(\frac{1 \text{ mole of } \text{H}_3\text{O}^+}{1 \text{ mole of } \text{HNO}_3} \right) = \frac{0.040 \text{ moles of } \text{H}_3\text{O}^+}{1 \text{ liter}}$$

$$= 0.040 \text{ M } \text{H}_3\text{O}^+$$

$$\text{pH} = -\log(\text{H}_3\text{O}^+) = -\log(0.040) = \boxed{1.40} \quad \checkmark$$

8. pH of 0.12 M of KOH?



$$\left(\frac{0.12 \text{ moles of KOH}}{1 \text{ liter}} \right) \left(\frac{1 \text{ mole of OH}^-}{1 \text{ mole of KOH}} \right) = \frac{0.12 \text{ moles of OH}^-}{1 \text{ liter}} = 0.12 \text{ M OH}^-$$

$$\frac{(\text{H}_3\text{O}^+)(0.12)}{0.12} = \frac{1 \times 10^{-14} \text{ M}^2}{0.12} \quad \left. \begin{array}{l} \text{pH} = -\log(8.3 \times 10^{-14} \text{ M}) \\ \text{pH} = \boxed{13.08} \end{array} \right\} \quad \checkmark$$
$$= 8.3 \times 10^{-14} \text{ M}$$

9. pH of 2.4 M of H_2CO_3 ? K_a of $\text{H}_2\text{CO}_3 = 4.3 \times 10^{-7}$
 K_a of $\text{HCO}_3^- = 7 \times 10^{-11}$



$$K_a = \frac{(\text{H}_3\text{O}^+)(\text{HCO}_3^-)}{(\text{H}_2\text{CO}_3)}$$

$$4.3 \times 10^{-7} = \frac{(x)(x)}{2.4} \Rightarrow 4.3 \times 10^{-7} = \frac{x^2}{2.4} \quad \checkmark$$
$$\sqrt{x^2} = \sqrt{1.032 \times 10^{-6}}$$

$$x = 1.015874 \times 10^{-3} = 0.0010 \quad \checkmark$$

$$\text{pH} = -\log(\text{H}_3\text{O}^+)$$

$$= -\log(0.0010) = \boxed{3.00} \quad \checkmark$$



11. K_a of $H_2PO_4^- = 6.3 \times 10^{-8}$

K_a of $H_2CO_3 = 4.3 \times 10^{-7}$

Compare strength

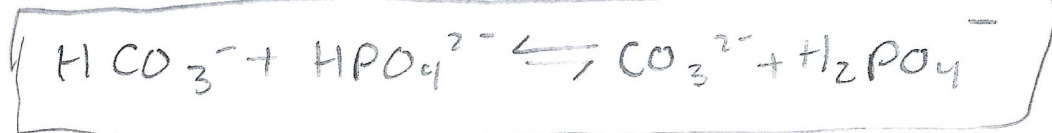
$K_a \cdot K_b = K_w$

stronger

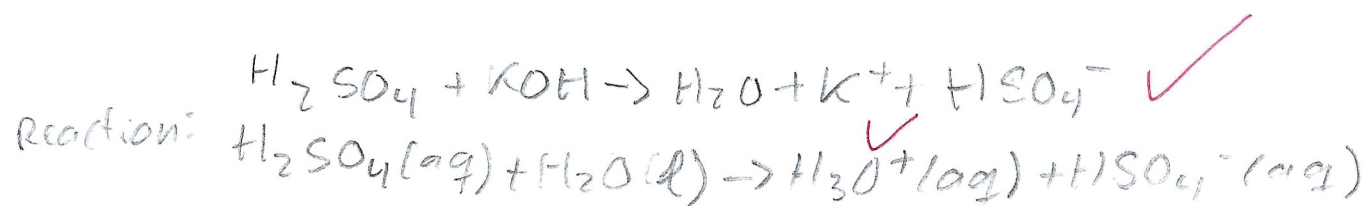
$$\frac{(6.3 \times 10^{-8}) K_b}{6.3 \times 10^{-8}} = \frac{1 \times 10^{-14}}{6.3 \times 10^{-8}}$$

$K_b = 1.6 \times 10^{-7}$

$$\frac{(\cancel{4.3 \times 10^{-7}}) K_b}{\cancel{4.3 \times 10^{-7}}} = \frac{1 \times 10^{-14}}{4.3 \times 10^{-7}} \quad K_b = 2.3 \times 10^{-8}$$

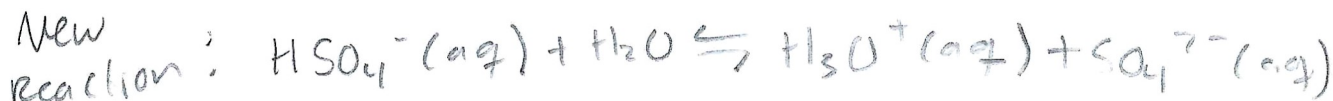


12. K_a of $HSO_4^- = 0.012$



$$\left(\frac{0.25 \text{ moles of } H_2SO_4}{1 \text{ liter}} \right) \left(\frac{1 \text{ mole of } H_3O^+}{1 \text{ mole of } H_2SO_4} \right) = \frac{0.25 \text{ moles of } H_3O^+}{1 \text{ liter}}$$

$= 0.25 M H_3O^+$



$$K_a = \frac{(H_3O^+)(SO_4^{2-})}{(HSO_4^-)} \Rightarrow 0.012 = \frac{(0.25+x)(x)}{(1.25-x)}$$

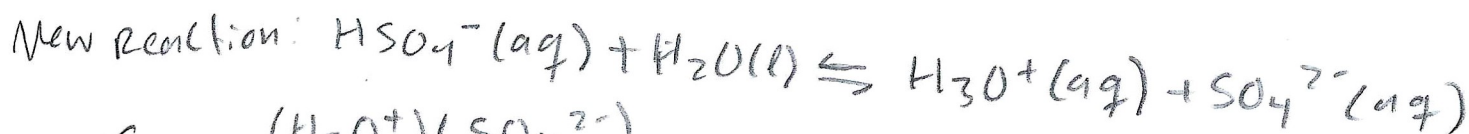
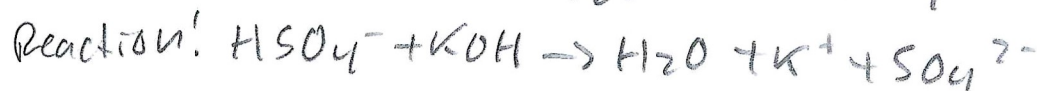
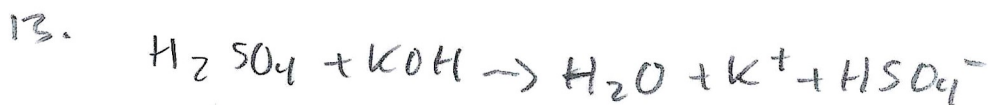


$$\frac{0.012}{(1.25-x)} = \frac{x^2 + 0.25x}{(1.25-x)} \cdot (1.25-x)$$

$$\begin{aligned} (1.25-x)(0.012) &= x^2 + 0.25x \\ 0.015 - 0.012x &= x^2 + 0.25x \\ -0.015 + 0.012x &+ 0.012x - 0.015 \\ \hline 0 &= x^2 + 0.26x - 0.015 \end{aligned}$$

$$\begin{aligned} x &= \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-0.26 \pm \sqrt{(0.26)^2 - 4(1)(-0.015)}}{2} \\ &= \frac{-0.26 \pm \sqrt{0.0676 + 0.06}}{2} = \frac{-0.26 \pm \sqrt{0.1276}}{2} \\ &= 0.050 \text{ or } -0.31 \end{aligned}$$

$$\text{pH} = -\log(\text{H}_3\text{O}^+) = -\log(0.30\text{M}) = \boxed{0.52} \quad \checkmark$$



$$K_4 = \frac{(\text{H}_3\text{O}^+)(\text{SO}_4^{2-})}{(\text{HSO}_4^-)} \Rightarrow 0.012 = \frac{x(0.25+x)}{(1.00-x)} \cdot (1.00-x)$$

$$\begin{aligned} 0.012(1.00-x) &= x^2 + 0.25x \quad \checkmark \\ 0.012 - 0.012x &= x^2 + 0.25x \\ -0.012 + 0.012x &+ 0.012x - 0.012 \\ \hline 0 &= x^2 + 0.26x - 0.012 \end{aligned}$$

$$\begin{aligned} x &= \frac{-0.26 \pm \sqrt{0.0676 - 4(1)(-0.012)}}{2} = \frac{-0.26 \pm \sqrt{0.1156}}{2} \\ &= \frac{-0.26 \pm 0.34}{2} = 0.04 \text{ or } -0.30 \end{aligned}$$

$$\text{pH} = -\log(0.04\text{M})$$

$$= \boxed{1.4} \quad \checkmark$$